IUPAC Periodic Table of the Isotopes Project

Learning Objectives

After having studied the materials associated with the IUPAC Periodic Table of the Isotopes Project, users should be able to do the following:

Explain that elements are made up of different types of atoms called isotopes and that most elements have more than one stable isotope.

Calculate the atomic weight of an element from the sum of the product of the abundance value and the atomic mass of each stable isotope of that element.

Explain that the isotopic abundances of many elements are not constant and vary in nature, as well as by the influence of accidental and deliberate human activities.

Explain why the standard atomic weight values of many elements are expressed as intervals.

Explain the causes of isotopic abundance variability.

Explain the difference between stable and radioactive isotopes and how properties of these isotopes and isotope ratios are related to different practical applications in everyday life.

Introduction to Isotopes

The Periodic Table of the Chemical Elements seen on chemistry classroom walls and the inside covers of chemistry textbooks provide properties of the known chemical elements in a progression from the lighter elements to the heavier elements.

Each of these elements are made up of atoms, while the atoms are composed of a central nucleus containing positively charged particles called protons and neutral particles called neutrons, which are slightly heavier than the protons. Around the outside of the nucleus are negatively charged particles called electrons, which are almost two thousand times lighter than either the protons or the neutrons. Normal atoms have an equal number of protons and electrons, so they are electrically neutral. If an atom is stripped of one or more of its electrons, it becomes positively charged and is called an ion of that element.

The various chemical elements are distinguished from one another by the number of protons in their nucleus, called the atomic number, Z. The lightest element, hydrogen, contains one proton and has atomic number, $Z = 1$. The heaviest of the naturally occurring chemical elements, uranium, contains ninety-two protons and has atomic number, $Z = 92$. There are a number of chemical elements, which do not occur in nature but these have been artificially produced, e.g., all of the elements that are heavier than uranium, from $Z = 93$ to $Z = 118$, have been produced synthetically. In addition, some elements that are lighter than uranium do not occur in nature.

All atoms of a given chemical element have the same atomic number, or number of protons in their nucleus. However, these atoms may have different numbers of neutrons in their nucleus. Since most of the mass of the atom comes from its protons and neutrons, the total mass is primarily made up from the number of protons and neutrons in the nucleus. The various masses of the atoms for a given element are called isotopes. If a chemical element, E, has P protons and N neutrons in the nucleus of one of its isotopes, this isotope is designated as E-x, where the mass number, x = P+N, is the total number of protons and neutrons. Only certain masses for a given element are energetically stable, i.e., they are constant with time. These particular masses are called the stable isotopes of that element. All of the other isotopes for that element are energetically unstable and are called unstable or radioactive isotopes (or radioisotopes for short). These radioactive isotopes will decay over time by emitting energy via particles or radiation, until they reach an energetically stable state, which will correspond to a stable isotope of that element or possibly of a nearby element. The time it takes for one half of the total number of atoms of that isotope to decay away is called the radioactive half-life of that isotope.

The Periodic Table was originally developed a century and a half ago because of the observed periodicity in the chemical and/or physical properties of the elements as these elements were listed in order of their increasing atomic weight (or relative atomic mass) values, which depends upon the isotopic masses of the different stable isotopes and the relative amounts of each stable isotope in the sample of that element. These atomic weight values were the original organizing principle of the Periodic Table and were considered to be constants of nature, similar to the speed of light. They were originally measured chemically relative to a scale standard of hydrogen = 1, or oxygen = 16. The oxygen scale was preferred because the heavy element values were close to whole numbers on this scale but not for the hydrogen scale. When isotopes were first discovered about a century ago, it was realized that measurements of the relative composition or amounts of the stable isotopes of a given element (called the isotopic abundance of that element or its mole fraction) combined with the atomic mass of these stable isotopes could provide an alternate method for determining an atomic weight value. Chemists used a scale where the element oxygen = 16, while physicists used a scale where the isotope, oxygen-16 = 16. When two other stable isotopes of oxygen were discovered in 1929, these two scales differed from one another. When the abundance of the oxygen isotopes in air and in water was found to be different, these two scales had a variable difference for the next twenty-five to thirty years. A solution to the problem was finally found in 1960/61, when everyone agreed to use the isotope, carbon-12 = 12 scale.

Some chemical elements have only a single stable isotope, while the majority of the elements have two or more stable isotopes. For elements with a single stable isotope, the isotopic mass is the same for all atoms of that element and it can be measured to better than one part in a million parts. The atomic weight is known with very high precision (many significant digits), since the isotopic abundance is exactly 100% (or has a mole fraction of one) and never varies. The atomic weight value of an element with only a single stable isotope can be considered to be a constant of nature.

For elements with two or more stable isotopes, the atomic weight depends on the relative amounts of each stable isotope. The isotopic ratio of these stable isotopes can vary in different samples of that element found in nature. This variation in the values of the stable isotope abundances can produce changes in the atomic weight value for that element. Thus, different samples of an element can have different atomic weight values. Although the atomic weight of an element in a specific sample may be known precisely, variations in the values from other samples of that element cause an uncertainty in the atomic weight that can be quoted to apply to all samples of that element. This type of uncertainty is not a result of poor measurements but a result of real variations over natural terrestrial samples of that element.

For many elements, the isotopic variation in nature of the atomic weight is much larger than the measurement uncertainty in the atomic weight value. This natural variation in the atomic weight value has often led users to mistake it for experimental uncertainty in the atomic weight. This problem has led to the introduction of intervals to express the upper and lower bounds of the atomic weight values, due to the natural variation of the isotopic abundance of stable isotopes in various samples of that element. To fully understand this effect, the IUPAC Periodic Table of the Isotopes has been prepared to help serve as an introduction to isotopes and as an aid in understanding the importance of isotopes in our everyday life. The human body contains the element potassium and one of its isotopes, potassium-40 (or ^{40}K), is radioactive. ^{40}K is part of the naturally occurring isotopic composition of potassium, so the human body is naturally radioactive, i.e., all of us are radioactive to a slight degree because of isotopes.

Variations in Isotopic Composition of the Elements

Stable isotopes of chemical elements generally occur in constant proportions. This accounts for the fact that atomic weight determinations on samples of a given element from widely different sources generally agree within experimental uncertainties. There are notable exceptions to this rule. The abundance of lead isotopes varies depending upon the source of ores, particularly those with uranium and thorium, because of the natural radioactive decay chain, as mentioned earlier. The isotope strontium-87 (or ^{87}Sr) has an abnormally high abundance in rocks that contain rubidium because rubidium-87 (or ^{87}Rb) is a naturally occurring β-emitter that will decay over time to ^{87}Sr. ^{87}Sr is often referred to as a radiogenic isotope because it can be a product of a radioactive decay. Radiogenic isotopes are some of the most important tools in geology (see the applications sections for information on radiometric dating and isotopic tracer).

The isotopic composition of elements in samples can vary both naturally and artificially (by human intervention). Isotopes of some elements can be separated from each other by physical, chemical and biological processes called isotopic fractionation. Many processes are mass dependent and readily observed in the case of the light elements, where the mass difference of the stable isotopes is significantly, e.g., hydrogen with mass 1 and 2. The percentage of deuterium, ^{2}H or D, is significantly different in warm water and cold water. Polar ice contains approximately ⅔ as much deuterium as ocean water because of the vapor pressure difference between H_2O and HDO causes isotopic fractionation, which depends on the temperature.

In the case of artificial differences in the isotopic composition, consider the enrichment of uranium. The lighter isotope, ^{235}U, is about ¾ of one percent of natural uranium but it is the isotope that undergoes nuclear fission with neutrons at an energy corresponding to room temperature. When an attempt was made to create a fission weapon (an atom bomb) during World War II, this light isotope was enriched to about 92% from just 0.72% by diffusion and electromagnetic separation.

The atomic weight of an element can be determined from the sum of the products of the isotopic abundance value and the isotopic mass of each stable isotope of an element. Let us consider a simple calculation of the atomic weight of carbon, with stable isotopes, ^{12}C and ^{13}C, whose isotopic abundance values will have simplified numbers of 99% and 1%, respectively. We will use simplified atomic mass values of 12 and 13 for these isotopes. The atomic weight for carbon would be $(0.99 \times 12) + (0.01 \times 13) = 12.01$ in this case. If another sample of carbon had simplified isotopic abundance values of 98% and 2%, the atomic weight would become $(0.98 \times 12) + (0.02 \times 13) = 12.02$ for this sample of carbon. Thus, a variation in the isotopic composition of an element in various samples in nature can result in a variation in the atomic weight value in those samples.

Isotopes of elements are important to us for a number of reasons. An obvious application already mentioned is the determination of the atomic weight of the element from stable isotopic abundance values.

Many applications rely on the properties of radioisotopes of the elements, such as their half-life and their decay properties. However, measured ratios of stable isotopes of some elements are also useful in everyday life. As will be shown later, the ratio, $^{13}C/^{12}C$, in the testosterone of the human body varies from that ratio in synthetic testosterone and can be used to determine doping in sports.

The Periodic Table of the Isotopes is color-coded to distinguish between elements which have no stable isotopes (colored white) and has no standard atomic weight, elements which have only one stable isotope (colored blue) and their standard atomic weight is a constant of nature, elements which have two or more stable isotopes or characteristic terrestrial isotopic composition (colored yellow) and their standard atomic weight is not a constant of nature and the measurement uncertainty exceeds any variation in their natural isotopic composition, and elements which have two or more stable isotopes (colored pink) and their atomic weight and isotopic composition vary in nature to the extent that this variation exceeds the measurement of their standard atomic weight. Atomic weight intervals are used to indicate the upper and lower bounds of the standard atomic weight values for these "pink" elements.

Nucleosynthesis

When you view the periodic table, a natural question arises of where did the chemical elements and their various isotopes come from? The process of creating these chemical elements from pre-existing protons and neutrons is called nucleo-synthesis. A simplified explanation of how nucleo-synthesis works requires some basic background information.

The atomic mass of an isotope that has Z protons and N neutrons, m (Z, N), is the sum of $Z \times m_p$ and $N \times m_n$ minus the binding energy of the isotope, where m_p and m_n are the atomic masses of the proton and the neutron, respectively, and the binding energy is the energy required to separate the isotope's nucleus into Z protons and N neutrons.

Albert Einstein's formula, $E = mc^2$, relates mass and energy, via the square of the speed of light, c^2. Small amounts of mass can be converted into large amounts of energy due to the large value of the speed of light, e.g., the decay of one pound of radium, ^{226}Ra, into radon, ^{222}Rn, and alpha particles, ^{4}He, from the reaction, $^{226}Ra \rightarrow {}^{222}Rn + {}^{4}He$, would produce the energy equivalent of almost 2/5th of a million pounds of TNT. A practical example of the conversion of mass into energy would be the fission of uranium fuel in a nuclear power reactor to generate heat, which converts water into steam and generates electricity. This is the mass into energy equivalent of the natural gas or coal burning power plants that burn gas or coal to generate heat to convert water into steam to produce electricity.

For each isotope, the binding energy per particle varies as the mass of each isotope increases. The binding energy initially increases as you move from the proton, ^{1}H, up to the iron isotope, ^{56}Fe, which has the peak binding energy per particle (this means that ^{56}Fe is the most tightly bound of all nuclei). Beyond ^{56}Fe, the binding energy per particle decreases. The result is that as you combine small nuclei to form a larger nucleus (in a process called nuclear fusion), there is energy released. Similarly, as you split a nucleus that is larger than ^{56}Fe into smaller nuclei (nuclear fission is an example of such a process) there is also energy released. This implies that for nuclear fusion to create any nuclei with masses above that of ^{56}Fe, you must add additional energy into the system for that to occur.

There are a number of types of nucleosynthesis, such as Big Bang nucleosythesis, stellar nucleosynthesis, explosive nucleosynthesis and cosmic ray spallation.

The Big Bang is thought to have occurred within the first three minutes of the beginning of the universe. It is responsible for the abundance of ^{1}H, ^{2}H, ^{3}He, ^{4}He and some ^{7}Be and ^{7}Li. The Big Bang occurred in a very short period until expansion and cooling stopped it. No heavy elements above beryllium were formed because the temperature and density was too small and there is no stable nucleus with eight nucleons in it.

After the Big Bang, nothing happened for a long period of time. Eventually, galaxy and star formation was produced by gravity and led to the synthesis of heavier elements during stellar nucleosynthesis. Matter coalesced to higher temperatures and densities. Reactions in stars involve charged particles, so stellar nucleosynthesis is a slow process. It involves the generation of elements from carbon to iron by nuclear fusion processes. The production of carbon from three alpha particles is a key factor because its lack of production was a problem for the Big Bang. Now heavier elements could be produced because various reactions could take place that released energy which kept the star from collapsing under gravitational forces.

With temperatures in the range of 10^8 to 10^9 degrees Kelvin and densities on the order of 10^5 to 10^8 grams per cm^2, the high temperatures provided energy for various reactions to become possible. Three alpha particles could produce ^{12}C; carbon and an alpha particle could produce ^{16}O; two carbon atoms could produce ^{20}Ne and an alpha particle; neon and an alpha particle could produce magnesium; Two oxygen atoms could produce a silicon atom; two silicon atoms could produce iron;
These reactions are the following:
The reactions to produce carbon are $^4He + {}^4He \rightarrow {}^8Be$ and $^8Be + {}^4He \rightarrow {}^{12}C$, which is called helium burning.
The reaction to produce oxygen is $^{12}C + {}^4He \rightarrow {}^{16}O$, also termed helium burning.
The reaction to produce neon is $^{12}C + {}^{12}C \rightarrow {}^{20}Ne + {}^4He$, which is called carbon burning.
The reaction to produce magnesium is $^{20}Ne + {}^4He \rightarrow {}^{24}Mg + \gamma$, which is called neon burning.
The reaction to produce silicon is $^{16}O + {}^{16}O \rightarrow {}^{28}Si + {}^4He$, which is called oxygen burning.
The reaction to produce iron is $^{28}Si + {}^{28}Si \rightarrow {}^{56}Ni + \gamma$, which is called silicon burning and is followed by two beta decays of $^{56}Ni \rightarrow {}^{56}Co \rightarrow {}^{56}Fe$.
These various fusion reactions released energy that kept the star from gravitational collapse.

As noted above, ^{56}Fe is the peak of the binding energy per particle curve, so no further nuclear fusion is possible above iron. As a result, stars cannot convert iron into any higher masses at this point via nuclear fusion. Other processes are required to create the heavier elements.

When fusion processes terminated and could no longer generate energy, electron capture reactions in iron and nickel isotopes destabilize the core, which then collapsed under the gravitational force and triggered a supernova explosion. There are two significant (α, n) reactions that have been taking place, the reaction $^{22}Ne + {}^4He \rightarrow {}^{25}Mg + n$ and the reaction $^{13}C + {}^4He \rightarrow {}^{16}O + n$. As a result of these reactions, there are free neutrons available. Two neutron capture processes take place on the iron and nickel nuclei that depend on the neutron densities available.

The term Nova (the Latin word for new) refers to a bright new star. A Supernova is a stellar explosion that is more energetic than a Nova. Explosive nucleosynthesis refers to nucleosynthesis that occurs in a core-collapse supernova. It involves a succession of rapid neutron capture reactions (called the r-process) on ^{56}Ni seed nuclei. These fast neutron captures continue until the nuclear force can no longer bind an extra neutron and a beta-decay occurs and a new chain of neutron capture reactions begin. The r process produces highly unstable nuclei, which have many neutrons and produces chemical elements up to uranium, plutonium and californium. This process is possible because the rate of neutron capture is faster than the beta-decay rate due to the high density of neutrons (on the order of 10^{22} neutrons/cm^3).

The other predominant mechanism for the production of heavy elements is the slow neutron capture reaction (called the s-process), which occurs at relatively low neutron densities (on the order of 10^7 neutrons/cm^3) and intermediate temperature conditions in stars. In the s-process, a stable isotope captures a neutron but the radioisotope produced is unstable and undergoes a beta-decay, when a neutron becomes a proton and the atomic number increases by one but the mass number A does not change. The radioisotope has decayed to its stable daughter before another neutron is captured. The final end point of the s process occurs at ^{209}Bi.

There is one other mechanism of rapid proton capture (called the rp process) which will account for the creation of additional isotopes in cosmic ray spallation reactions. In this process, fast protons in cosmic rays react with stellar material.

There are some three thousand unstable isotopes discovered so far. There are many thousands more unstable isotopes that have not yet been discovered.

To fully explain the origin of each isotope of the chemical elements require a more detailed understanding of the physics of nuclei and nuclear reactions, which is beyond the level of this introductory material.

Applications of Stable and Radioactive Isotopes

There are many techniques involved in the use of both stable and radioactive isotopes. The following information will indicate some of the techniques that have been used to apply isotopes or isotope ratios to help solve problems in a variety of scientific fields.

It should be noted that radioisotopes undergo a process called radioactive decay to a stable isotope by emitting alpha particles (helium nuclei), beta particles or positrons (negatively charged or positively charged electrons) and also radiation (gamma rays). Half-life is the time it takes for one half of all of the atoms of a given radioisotope to decay. We use a value of 10^{10} years (or 10 billion years) as the dividing line between a stable isotope and a radioisotope. An isotope with a half-life equal to or greater than ten billion years is considered to be stable in this discussion.

Geochronology – is the science of dating and determining the time sequence of events in the history of the earth. Numeric dating (absolute dating) involves determining geological ages of a sample, such as a fossil, rock or geologic event in units of time, usually years. This type of dating uses radiometric or isotopic methods which samples radiogenic elements and their by-products of radioactive decay. Knowing the relative amounts of radioactive (parent) and radiogenic (daughter) isotopes with a known half-life, one can calculate the elapsed time since the product was formed. Some examples of the use of this method are the following: potassium-40/argon-40 geochronology; argon-40/argon-39 geochronology; uranium series (or uranium-lead geochronology) methods; lead-210 geochronology; radiocarbon geochronology; rubidium-87/strontium-87 geochronology; samarium-147/neodymium-143 geochronology; tritium/helium-3 geochronology; rhenium-187/osmium-187 geochronology; uranium-234/thorium-230 geochronology;

Isotope Fractionation – It was mentioned earlier that the mass of an atom depends upon the number of protons and neutrons in the nucleus. For a chemical element with a given atomic number, the number of protons in each atom is fixed, so the mass of its isotopes depend on the number of neutrons. The mass difference between the isotopes will result in a partial separation of the light isotopes from the heavy isotopes during various chemical reactions and physical processes such as diffusion and vaporization and is called mass dependent isotope fractionation. The effect of fractionation is most noticeable for the light elements, where the mass difference between the isotopes would be greatest.

Isotopic Labeling – is a technique for tracking the passage of a sample of a material through a system. A material is labeled with an isotope which would be unusual in its chemical composition. If these unusual isotopes are eventually detected in another part of the system, they must have come from the labeled material. If the unusual isotope has a larger or smaller concentration than the element's normal isotopic composition, this difference can be detected by mass spectrometry. If the unusual isotope has a different mass than the stable isotopes and is radioactive, it can be detected by the radiation that it emits.

Radioisotope usage in the medical field – nuclear medicine uses radioisotopes to provide diagnostic information about the functioning of specific human organs or to treat them. Rapidly dividing cells are particularly sensitive to damage by radiation. Some cancerous growths can be controlled or eliminated by irradiating the area. This is radio-therapy. Radiotherapy can be used to treat some medical conditions, especially cancer, using radioisotopes to weaken or to destroy particular targeted cells.

Computed tomography (CT) or computed axial tomography (CAT) scanning is a medical imaging procedure that uses x-rays to show cross sectional images of the body.

Positron emission tomography (PET) scans detect pairs of gamma-rays emitted indirectly by a positron emitting radioisotope (tracer). Images of these tracer concentrations in 4-dimensional space (time is the fourth dimension) within the body are reconstructed by computer analysis, often with the aid of a CT x-ray scan performed on the patient during the same session in the same machine.

Some therapeutic procedures are palliative to relieve pain, such as cancer induced bone pain. Radiotherapy may be preferable to traditional pain killers such as morphine because it improves patients' quality of life, allowing them to be more lucid during time spent with family.

Radioactive products which are used in medicine are referred to as radiopharmaceuticals. Every organ in our body acts differently from a chemical point of view. Some chemicals preferentially absorbed by specific organs are called targeting agents, such as iodine in the thyroid, strontium in bone and glucose in the brain. When a radioactive form of one of these substances enters the body, it is incorporated into normal biological processes and excreted in the usual ways and its process can be tracked. Radiopharmaceuticals can be used to examine blood flow to the brain, to evaluate functioning of the liver, heart or kidneys to assess bone growth and to predict the effects of surgery and assess changes since a treatment has begun. In sports medicine, radiopharmaceuticals can be used to diagnose stress fractures, which are not generally visible in x-rays.

Isotopes in industry – Radioisotopes are used in industry in many ways, such as in radiography, in gauging applications, in mineral analysis, in flow tracing, in gamma sterilization for medical supplies and for food preservation. Gamma radiography is used to scan luggage at airports. Gamma rays show flaws in metal castings or welded joints. Critical components can be inspected for internal defects but without damaging the component or making it radioactive. Unlike ×-rays, radioactive sources are small and do not require power, so they can be transported easily to remote areas, where there is no power.

Radioisotope gauging uses the fact that radiation will be reduced in intensity by matter located between the radioisotope and a detector. The amount of reduction can be used to gauge the presence or absence of the material or even to measure the quantity of material between the source and the detector. An advantage of this method is no contact with the material being measured. Radioisotopes are used to analyze the contents of mineral samples, such as the continuous measurement of mineral slurry density in processing the liquids used to separate a metal from its ore. Gamma-ray transmission or scattering can determine the water content of coal on conveyer belts. The gamma-ray interaction varies with the atomic number (number of electrons) of the atom, hydrogen and oxygen versus carbon. Material coatings can be measured because some radiation is scattered back toward the radiation source and analysis of the back-scattered radiation provides information on the material’s coating.

Radioactive tracers can be used to trace small leaks in complex systems such as power station heat exchangers. Flow rates of liquids and gases in pipelines, as well as large rivers, can be measured accurately with the assistance of radioisotopes.

Smoke detectors are the most numerous of all devices that employ radioactive isotopes worldwide. Smoke detectors use the radioisotope, ^{241}Am, to ionize atoms of air (knock out external electrons from the atom) producing small electrical current that is monitors. When smoke or steam enters the ionization chamber, it disrupts the current. The smoke detector senses the drop in current between the plates and sets off the alarm.

Glossary of Terms

Atomic Mass – is the mass value of atomic species in unified atomic mass units on a scale in which an atom of carbon-12 is defined as equal to 12.0 exactly.

Atomic Number – of a chemical element is equal to the number of protons in the nucleus of that element.

Atomic Weight – (relative atomic mass) of an element in a given sample is equal to the mean value of the atomic masses of all of the individual stable atoms of that element in the sample.

Atomic Weight Interval – is the upper and lower bounds determined for the standard atomic weight of elements, when its isotopic and atomic weight variation in nature exceeds the measurement uncertainty of its standard atomic weight.

Atomic Weight Uncertainty – is the measurement uncertainty of the standard atomic weight of a given chemical element.

Half-life (radioactive) – is the time interval that it takes for the total number of atoms of any radioactive isotope to decay and leave only one-half of the original number of atoms.

Isotopic Abundance – (or mole fraction) is the percentage of all stable isotopes of a given chemical element made up of a given mass number.

Isotope – is one of two or more species of atoms of a given element (having the same number of protons in the nucleus) with different atomic masses (numbers of neutrons in the nucleus).

Isotope Number – (or mass number) is the total number of protons and neutrons in a given atom of an element.

Isotopic Labeling – is a technique for tracking the passage of a sample of a substance through a system. The substance is "labeled by including unusual isotopes in its chemical composition. If these unusual isotopes are later detected in a certain part of the system, they must have come from the labeled substance.

Mole Fraction – (or isotopic abundance value) is the percentage of all stable isotopes of a given chemical element made up of a given mass number.

Nucleosynthesis – is the process of creating chemical elements from pre-existing protons and neutrons.

Periodic Table of the Chemical Elements – is a table of the elements beginning with atomic number one and listing all of the known chemical elements and providing various chemical properties of these elements.

Periodic Table of the Isotopes – is a table of the chemical elements beginning with atomic number one and listing all of the known chemical elements and providing diagrams of the isotopic composition of all of the stable isotopes of each element. The interactive version of this table also provides various applications of stable isotope ratios and radioactive isotopic applications that have proven useful in everyday life.

Radioactive Decay – the process by which unstable (or radioactive) isotopes lose energy by emitting alpha particles (helium nuclei), beta particles (positive or negative electrons), gamma radiation, neutrons or protons to reach a final stable state.

Radioactive Isotope – is an isotope which is not stable and spontaneously disintegrates by undergoing the radioactive decay process in a time period determined by its radioactive half-life.

Radioactive Tracer – (radioactive label) is a substance containing a radioisotope that is used to measure the speed of chemical processes and to track the movement of a substance through a natural system such as a cell or tissue.

Radiogenic Isotope – is an isotope that is produced by the process of radioactive decay.

Radio-isotopic labeling – is a special case of isotopic labeling, where the substance is labeled by using radioactive isotopes in its chemical composition.

Stable Isotope – is an isotope which remains unchanged over time and does not undergo the process of radioactive decay.

Standard Atomic Weight Value and its Associated Uncertainty – are evaluated quantities assigned by the Commission on Isotopic Abundances and Atomic Weights to encompass the range of possible atomic weights of a chemical element that might be encountered in all samples of normal terrestrial materials. The uncertainty on the Standard Atomic Weight of an element is based on either the measurement uncertainty of the atomic weight in a representative sample or the interval over which known variations in atomic weights among normal terrestrial materials have been reported, whichever is larger.